3 Chemical Equilibrium: Le Chatelier's Principle

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- **Observe several different chemical reactions that illustrate chemical equilibrium.**
- **Observe changes that illustrate the Le Chatelier principle.**

DISCUSSION

In this experiment you will see changes in color, and you will watch precipitates form or disappear. Since each of these changes will occur because an equilibrium shifts, you will be asked to explain your observations according to Le Chatelier's principle. For an even better learning experience, however, apply your knowledge of Le Chatelier's principle to the given chemical system *before* you do each procedure. Predict what you think ought to occur, and then verify your understanding by direct experiment.

Volumes of solutions used in this experiment generally do not need to be measured to great accuracy. Before beginning the experiment, obtain 1 mL of water in your small graduated cylinder, then pour it into a small test tube and notice its height. From that, you should be able to estimate 0.5 mL, 1 mL 1.5 mL, etc. Remember that a small test tube holds 10 mL when filled to the top. You can also quickly estimate how many drops of solution from a small dropper bottle approximately equal 0.5 mL, 1.0 mL, etc. Estimating volumes in this way can save a great deal of time, if you are reasonably careful.

PROCEDURE

A. Equilibrium in Solution

The hexathiocyanatoferrate(III) complex ion. Complex ions form when certain ions combine with other ions or molecules. The iron(III) ion and the thiocyanate ion form a complex ion according to the following equilibrium equation.

$$Fe^{3+}{}_{(aq)} + 6SCN^{-}{}_{(aq)} \leftrightarrows Fe(SCN)_6{}^{3-}{}_{(aq)}$$

This complex ion has a deep, blood-red color.

1. Add 0.5 *mL* of 0.1 *M* $FeCl_3$*(aq)* to 0.5 *mL* of 0.1 *M* KSCN*(aq)*, and then add 15 *mL* of distilled water. To a 2-mL portion of this mixture, add 1 mL of 0.1 *M* KSCN*(aq)*. Observe any difference in the color from the original solution and explain your observations.

Because you are combining equal volumes of equal concentrations, and because they combine in a 1 to 6 mole ratio, Fe^{3+}(aq) should be in excess. After you establish the initial equilibrium, and after dilution with 15 mL of water, when you add more KSCN(*aq*) you should see the results of a shift in the equilibrium. If you see no change, something went wrong and you should try again.

2. Determine how much distilled water must be added to 0.5 mL of 0.1 *M* $FeCl_3$*(aq)* in order to make it 0.001 *M*. Then carry out the dilution.

3. Add 0.5 mL of this 0.001 *M* $FeCl_3$*(aq)* to 0.5 mL of 0.1 *M* KSCN*(aq)*, and then add 15 mL of distilled water. To a 2 mL portion of this mixture, add 1 mL of 0.1 *M* $FeCl_3$*(aq)*. Observe any change in color and explain your observations. In this case, SCN^-(aq) should be in excess.

After you establish the initial equilibrium, and after dilution with 15 mL of water, when you add more $FeCl_3$(aq) you should see the results of a shift in the equilibrium. If you see no change, something went wrong and you should try again.

The tetrachlorocuprate(II) complex ion. In a water solution, the copper(II) ion is bonded

- **CAUTION: Compounds of either antimony or barium are poisonous.**
- **DO NOT pour solutions of heavy metal compounds down the drain.**

to four water molecules to form the hydrated complex ion, $Cu(H_2O)_4^{2+}$. This ion has the characteristic blue color that is normally associated with copper solutions. High concentrations of chloride ions replace the water and form a different complex ion that has a green color, as the following equation illustrates.

$$\underset{\textit{blue}}{Cu(H_2O)_4^{2+}{}_{(aq)}} + 4Cl^-_{(aq)} \leftrightarrows \underset{\textit{green}}{CuCl_4^{2-}{}_{(aq)}} + 4H_2O_{(\ell)}$$

4. To 1 mL of 0.1 *M* $Cu(NO_3)_2$*(aq)*, add 1 *mL* of concentrated (12 *M*) HCl*(aq)*. Note any color change, and explain what happened.

5. Now add 5 mL of distilled water to the mixture. Explain any new change in terms of the Le Chatelier principle. Add more water if necessary until you see a color change. The solution will become dilute enough that the color may be hard to see. The best way to see it is to look down the length of the test tube from the top. Discard the solution into the waste container.

The chromate-dichromate equilibrium. The yellow chromate ion will react with the hydronium ion from an acid to form the orange dichromate ion according to the following equation. Chromate ion is yellow in solution, while dichromate ion is orange. Carefully note the difference between these two colors as you do this part of the experiment.

$$\underset{\textit{yellow}}{2CrO_4^{2-}{}_{(aq)}} + 2H_3O^+_{(aq)} \leftrightarrows \underset{\textit{orange}}{Cr_2O_7^{2-}{}_{(aq)}} + 3H_2O_{(\ell)}$$

6. Place 1 mL of 1.0 *M* K_2CrO_4*(aq)* in a test tube. Add two drops of 3.0 *M* H_2SO_4*(aq)* to the test tube. Mix the two and note the change. Because there is some hydronium ion in water, the equilibrium is already established, even in a solution of potassium chromate. Add more sulfuric acid, if necessary, until you see a distinct color change.

7. Now add a few drops of 6 *M* NaOH*(aq)* until you get a second change. (Add enough sodium hydroxide to produce a distinct color change.)

8. Add more acid, and see if you can get the first change again. Then add more base, and see if you can get the second change again. Explain these changes in terms of Le Chatelier's principle.

Discard the solution into the waste container.

B. Equilibrium in Saturated Solutions

Saturated sodium chloride. When a precipitate forms in a solution, the solid is in equilibrium with the ions that remain dissolved. This can be illustrated with an equilibrium equation. For example, a saturated solution of sodium chloride is described by

$$NaCl(s) \leftrightarrows Na^+(aq) + Cl^-(aq)$$

This equation describes the equilibrium in a saturated solution, even if you don't see any solid present.

9. Add 1 mL of 12 *M* (concentrated) HCl*(aq)* to 5 mL of 5.4 *M* (saturated) NaCl*(aq)* solution. Note the change, and write the equilibrium equation for a saturated NaCl solution. Explain your observations in terms of the Le Chatelier principle. **PLEASE be careful with 12 *M* HCl**. It is not user-friendly. Add more than 1 mL, if necessary, until you see a change.

10. Now add 5 *mL* of distilled water to the mixture. Account for any change. Add more water, if necessary, and stir the mixture until you see a change.

Antimony(III) oxychloride. Antimony(III) chloride, $SbCl_3$, is a soluble compound that reacts with water to form insoluble antimony(III) oxychloride, SbOCl, according to the following equation:

$$SbCl_3(aq) + H_2O(l) \leftrightarrows SbOCl(s) + 2\ HCl(aq)$$

11. Obtain 0.5 mL of a solution which is 0.1 *M* $SbCl_3$*(aq)* and 6 *M* HCl*(aq).* Why does this solution contain no visible precipitate of SbOCl?

12. Add 5 mL of distilled water. Record your observation, and explain it. Add more water, if necessary, until you see a change.

13. Now add 1 mL of 12 *M* HCl*(aq)*, and account for the result. Add more 12 *M* HCl, if necessary, and stir the mixture until you see a change. Dispose of the solution in the specially marked container.

Barium chromate. Barium chromate is insoluble, whereas barium dichromate is soluble.

14. To a 1 mL sample of 0.1 *M* $BaCl_2$*(aq)*, add two drops of 1.0 *M* K_2CrO_4*(aq)*. Note the reaction, and write the *net-ionic* equation for the reaction that comes to equilibrium.

15. Now add 1 mL of 6 *M* HCl*(aq)*, and observe the changes. Add more HCl, if necessary, until you see two changes—one of which is a color change. Explain the result in terms of the chromate-dichromate equilibrium in Part A.

Whenever two equilibria occur in the same mixture as in this case (chromate ion is in equilibrium with the solid barium chromate and with the dichromate ion), *both* equilibria must be satisfied. Dispose of the solution in the waste container.

3 Chemical Equilibrium

Name ______________________________
Section __________________Locker ____
Instructor ___________________________

A. Equilibrium in Solution

The hexathiocyanatoferrate (lll) complex ion.

1. Describe the change you observed when you added 1 mL of 0.1 M KSCN to the 2 mL portion of the diluted solution. Copy the equation from the procedure and explain your observations in terms of LeChatelier's Principle.

2. Show your calculations to determine the final volume of the 0.001 *M* $FeCl_3$ solution.

3. Describe the change you observed when you added 1 mL of 0.1 M $FeCl_3$ to the diluted mixture of 0.001 M $FeCl_3$(aq) and 0.1 M KSCN(aq). Refer to the equation in step 1 and explain your observations in terms of LeChatelier's Principle.

The tetrachlorocuprate (II) complex ion.

4. Tell what happened when you added HCl(aq) to the copper nitrate solution. Write down the equation for the reaction.

5. Describe the change, and explain what occurred when you added the water (in terms of LeChatelier's Principle).

The chromate-dichromate equilibrium

6. Describe the change, copy the equation, and explain what occurred when you added the sulfuric acid to the potassium chromate solution.

7. What was the observed change when you added sodium hydroxide to the solution? Neither the Na^+ ion nor the OH^- ion appears in the chromate-dichromate equilibrium. Why, then, did the equilibrium shift?

8. Explain the color changes in terms of Le Chatelier's principle.

B. Equilibrium in Saturated Solutions

Saturated sodium chloride.

9. What did you observe when you added concentrated hyrochloric acid to saturated sodium chloride? Write the equation for the equilibrium, and explain what you observed in terms of LeChatelier's Principle.

10. Describe the effect water had on the mixture. Explain the change properly in terms of the LeChatelier principle.

Antimony(III) oxychloride.

11. The equilibrium equation shows that $SbCl_3$ reacts with water to form insoluble SbOCl. Why does the solution of antimony(III) chloride have no visible precipitate in it?

12. When you added water to the solution, what happened? Copy the equation, and explain your observation.

13. What effect did the concentrated hydrochloric acid have? Why?

Barium chromate.

14. Write the net-ionic equation for the reaction you observed between barium chloride and potassium chromate. Indicate the states of all reactants and products.

15. What happened when HCl(*aq*) was added? Record both changes that you observed. To explain these observations, first tell what effect the addition of H_2SO_4 had on the chromate-dichromate equilibrium (see Step 6). Then explain how that equilibrium shift in turn affected the barium chromate equilibrium (see Step 14). Write both equations again here as part of your explanation.

APPLICATION OF PRINCIPLES

1. The silver ion, Ag^+, forms the colorless diamminesilver(I) complex ion, $Ag(NH_3)_2^+$, that is soluble, when it is in an ammonia solution. If ammonia is added to a solution that contains a AgCl precipitate, the solid dissolves completely. Write a net-ionic equation for the equilibrium involved, and explain the shift that takes place.

2. When concentrated H_2SO_4 (18*M*) is added to a saturated solution of sodium sulfate, a white precipitate is formed. Write a net-ionic equation for the equilibrium in saturated sodium sulfate, and explain the change in terms of Le Chatelier's principle.

3. A solution contains the blood-red hexathiocyanatoferrate(III) complex ion, $Fe(SCN)_6^{3-}$, in equilibrium with the iron(III) ion, Fe^{3+}, and six thiocyanate ions, SCN^-. If the solution is diluted with water, will the color deepen, fade, or stay the same? Why? In addition to predicting whether the solution will deepen, fade, or stay the same, give TWO reasons for that prediction. One of those reasons will involve Le Chatelier's Principle, and one will not. (At least, that will be the case if you make the right prediction!)